

Course Review of AP Chem

ATOMIC STRUCTURE

1. Aufbau Principle
2. Exception to rule: no nd^4 or nd^9 for transition atoms. Can have this configuration for ions.
3. Paramagnetic and diamagnetic
4. For transition ions remove np and ns electrons first then remove $(n-1)d$ electrons.
5. Atomic orbital defined by the first three quantum numbers(n , l , and m_l)
 - A) n defines distance from nucleus, l defines shape of probability diagram and m_l defines orientation of probability diagram.
6. Quantum numbers(n , l , m_l and m_s)
7. Shapes of probability diagrams (s , p and d)
8. Electron dot configuration
9. Energy, wavelength and frequency of a photon
10. Excited state and forbidden state ($2d$)
11. Line spectrum(emission and absorbance)
12. Hund's rule and Pauli's Exclusion Principle
13. Heisenberg Uncertainty Principle
14. Rutherford gold foil experiment

PERIODIC RELATIONSHIPS

1. Atomic and ionic radius
 - A) Across row(period) Noble gas core(Group 18) constant and more protons in nucleus therefore attraction force on valence electrons increase and radius decreases. The number of filled quantized energy levels are the same.
 - B) Down a column(family) The number of filled quantized energy levels increases therefore less attraction force on valence electrons and the atomic radius increases.
 - C) Comparing isoelectronic ions Na^+ and Ca^{++} The proton/electron ratio is larger in calcium ion which increases the attraction force between the nucleus and the remaining electrons. Ca^{++} has smaller ionic radius.
 - D) Comparing isoelectronic negative ions O^{2-} and N^{3-} The proton/electron ratio is smaller in the nitride ion causing less attraction forces on the remaining electrons. N^{3-} has a larger ionic radius.
2. Ionization Potential(energy)
 - A) Across row(period) Stronger attractions forces on valence electrons due to smaller size of atom. Ionization energy generally increases.
 - 1) EXCEPTIONS: Column 2 and 13 Column 13 has a lower first IE because the np electron is further from the nucleus.
 - Column 15 and 16 Column 16 has a lower first IE because its last electron fills an atomic orbital causing more repulsion therefore easier to remove.
 - B) Down a column(family) First ionization energy decreases because the atomic radius increases and there is less attraction forces between the nucleus and the valence electrons.
 - C) Comparing first and second ionization energies Column 1 Second ionization energy is much greater than first because you are removing an electron from a filled quantized energy level that is closer to the nucleus.
 - D) Comparing first and second ionization energies Column 1 and 2 The second ionization energy of column 1 is much great than the second ionization energy of column 2 because you are removing an electron from a filled quantized energy level that is closer to the nucleus in column 1 and removing a valence electron in column 2.
3. Electron Affinity
 - A) $X(g) + e^- \rightarrow X^-(g)$ Energy change associated with the addition of an electron to a gaseous atom.

CRYSTALS

1. MOLECULAR

A) Lewis configurations

- 1) To have a pi bond must have two elements in the molecule that have two or more unpaired electrons in the electron dot configuration.
- 2) Account for all valence electrons.
- 3) Resonance: delocalized pi bond spread over several atoms(usually oxygen and central atom).
- 4) Pi bonds are perpendicular to sigma bonds.
- 5) Add electrons for negative ions and subtract electrons for positive ions.
- 6) Group 13 as a central atom usually has six electrons around it. Good Lewis acid(electron pair acceptor).
- 7) Carbon, nitrogen and oxygen as a central atom can only have eight electrons around it. There are no 2d orbitals in the second energy level.

B) Hybridization

- 1) Only lone pair electrons or unpaired electron and sigma bonding electrons go in hybrid orbitals. Pi bonding electrons ARE NOT in hybrid orbitals.
- 2) Only sp and sp² hybrid orbitals have the possibility of having pi bonds. Pi bonds exist only in pure p orbitals which are perpendicular to the hybrid orbitals and that is why pi bonds are weaker in strength than sigma bonds.
- 3) Need to know shapes associated with the hybrid orbitals

C) VSEPR

- 1) Must give electron pair geometry and molecular geometry(sigma bonds)
- 2) Electron pair geometry
 - a) 2 pairs linear 180°
 - b) 3 pairs trigonal planar 120°
 - c) 4 pairs tetrahedral 109.5°
 - d) 5 pairs trigonal bipyramidal 120°, 90° and 180°
 - e) 6 pairs octahedral 90° and 180°
- 3) Non-dipole molecules: identical polar bonds cancel out due to molecular shape.
- 4) Dipole molecules: polar bonds do not cancel out due to molecular shape. If you have a symmetrical molecule could have a dipole due to different polarities of the bonds.

D) Intermolecular Forces (IMF)

- 1) London dispersion forces all molecules have this IMF due to random movement of electrons. Non-dipoles only have this force. More electrons in a non-dipole, the stronger the dispersion force.
- 2) Dipole-dipole forces between dipoles
- 3) Hydrogen bonding occurs when a molecule has hydrogen bonded to oxygen or nitrogen or fluorine. The hydrogen electron is slightly displaced away from the bare proton of hydrogen causing a large polarity on the hydrogen atom.

E) Crystal lattice

- 1) Molecules are located at the lattice site in the unit cell and the molecules are held together by intermolecular forces.

F) Solubility

- 1) Dipoles will dissolve in dipoles and non-dipoles will dissolve in non-dipoles.

2. Ionic

A) Positive and negative charge density

- 1) The smaller the ionic radius, more concentrated the positive or negative charge and a greater attraction force for the negative or positive ion which results in a stronger ionic bond and a higher melting point.

B) Crystal lattice

- 1) At the lattice sites in the unit cell there are positive and negative ions held together by the Coulombic attraction force between opposite charged ions.

3. Metallic

A) Crystal lattice

- 1) At the lattice sites in the unit cell there are positive ions held together by mobile valence electrons traveling through empty valence orbitals.

B) Positive charge density

- 1) The smaller ionic radius will have the greater concentration of positive charge therefore having a stronger hold on the valence electrons. This will result in a higher melting point.
- 2) Example: The melting points of Column 1 (alkali metals) decrease going down the column.

4. Covalent network crystals

A) Crystal lattice

- 1) At the lattice sites in the unit cell there are atoms held together by strong covalent bonds. This crystal has the highest melting points.
- 2) Examples: Carbon in the form of diamonds. The carbon atoms are arranged in a tetrahedral structure, sp^3 hybrid. Carbon in the form of graphite. The carbon atoms are arranged in a trigonal planar, sp^2 , configuration and the layers are held together by delocalized pi bonds. Good conductor of electricity and a good lubricant. Silicon dioxide, made component in glass and sand. Silicon and CSi (cover for drills)

GASES

1. Ideal Gas

- A) No intermolecular forces between gas molecules due to high kinetic energy content of the molecules.
- B) The actual volume of the gas particles is insignificant when compared to the volume of the container.
- C) Molar volume(one mole) of an ideal gas at STP($T = 273$, $P = 1$ atm) is 22.4 liters.

2. Non-ideal Gas

- A) At the boiling point and just above the boiling point, intermolecular forces can exist resulting in a lower pressure than the ideal pressure. Since the molecules are sticking together due to IMF, the number of moles of gas particles is smaller. (low temperature)
- B) At very high pressure conditions, the volume of the gas particles must be taken into consideration.
Actual volume of gas particles =
 $V_{\text{container}} - V_{\text{gas particles}}$

3. Kinetic Energy

- A) $KE = \frac{1}{2} (\text{mass}) (\text{average velocity})^2$
- B) Kinetic energy is associated with Absolute temperature (Kelvin). When the Kelvin temperature is doubled the velocity of the gas particles increase by a square root of 2. Look at equation. Must increase temperature four fold to double velocity.

4. Diffusion and effusion

- A) The larger molar mass will have a smaller velocity.(square root of molar mass) Graham's law
- B) Direct relationship with time and square root of molar mass, indirect relationship between distance traveled and square root of molar mass.

5. Gas Stoichiometry

- A) Watch out for limiting reactant and finding total pressure when you have more than one gas involved. Use summation of gas moles including excess reactant.
- B) When gas is collected over water, must subtract water vapor pressure(nonideal gas) before placing pressure(dry gas) into ideal gas law.

SOLUTIONS

1. Concentration

- A) mole fraction
- B) % by mass
- C) molarity
- D) molality
- E) need density of solution to get molarity
- F) Dilution $M_c V_c = M_d V_d$ d is dilute and c is concentration

2. Vapor Pressure

- A) Temperature dependent only
- B) The vapor pressure of a solution (colligative properties)
 - 1) If you have a nonvolatile solute, the vapor pressure of the solution will depend on how many moles of solute particles are in solution. The lowering of the solution vapor pressure is due to the concentration of the solute. If the solute is a molecular compound then $i = 1$. If the solute is a soluble ionic compound(an electrolyte) the i will be greater than one. $P_{\text{soln}} = X_{\text{solvent}} P^{\circ}_{\text{pure solvent}}$ or change in temperature = $K m i$
 - 2) If you have a volatile solute, the vapor pressure of the solution will depend on the mole fraction(X) of the solute and solvent. Must find pressure of each and add the pressures. (Raoult's law) $\text{Vapor pressure of solution} = X_{\text{solvent}} P^{\circ}_{\text{pure solvent}} + X_{\text{solute}} P^{\circ}_{\text{pure solute}}$
 - 3) Find molar mass by using freezing point depression and boiling point elevation. The constant(K) is based on the solvent.

3. Solubility "like dissolves like"

ELECTROCHEMISTRY

1. Galvanic cell (spontaneous reaction)

- A) Half-reactions and overall reaction
- B) Finding the standard potential, E° using reduction potential table. $Q = 1$ and all concentrations are 1.0 M and partial pressures of gases are 1.0 atm.
- C) Label anode(oxidation) and cathode(reduction)
- D) Label flow of electrons (anode to cathode)
- E) Label flow of anion and cation in salt bridge and why?
- F) Predict the potential of a cell not at standard conditions using the Nernst equation and Q.
- G) At equilibrium Q becomes K in the Nernst equation and the potential of the cell is zero ($E = 0$).
- H) A reaction is spontaneous if the $E > 0$, $\Delta G < 0$ and K is a very large number.
- I) Find ΔG using the potential of the cell. Answer is large and the unit is joules because a volt is joules/coulomb
- J) Find the number of grams plated out at the cathode or volume of gas at one of the electrodes using current (ampere) and time (seconds). $Q = IT$

2. Electrolysis (non-spontaneous reaction)

- A) Half-reactions and overall reaction
- B) Find the number of grams plated out at the cathode or volume of gas at one of the electrodes using current(ampere) and time(seconds).
- C) In order for electrolysis to occur, must have an external energy source to drive the reaction. This reaction is non-spontaneous. The external energy source dictates the cathode and anode. The electrons flow into the cathode just like the galvanic cell.

KINETICS

1. Rate is always a positive value.
 - A) The rate of disappearance and rate of formation can be found by using the balance chemical equation provided one rate is known. Usually the rate is in the data table.
2. Rate law expression
 - A) Three ways to find rate law expression.
 - 1) Find the rate law expression by using experimental data comparing the rates of two experiments. Math
 - 2) Find the rate law expression by graphing. Linear function: Concentration vs time (zero order) \ln concentration vs time (first order) and $1/\text{concentration}$ vs time (second order) Should know what the slope on the linear function represents. If you get a curve when plotting Concentration vs time then the tangent to the curve is negative rate.
 - 3) From the mechanism use the coefficients in the rate determining step (slow step) for the order of the reaction for each reactant. This gives you the rate law expression. Can do because the rate determining step is an elementary step. Since each step in a mechanism is an elementary step, the coefficients for the reactants becomes the order of reaction for each reactant in the rate law expression.
3. Rate constant
 - A) The unit for the rate constant will vary depending on the total order of the reaction. Zero order mole $\text{L}^{-1}\text{time}^{-1}$ First order $1/\text{time}$ Second order $\text{L mole}^{-1} \text{time}^{-1}$ Third order $\text{L}^2\text{mole}^{-2}\text{time}^{-1}$ Don't need to memorize.
4. Reaction mechanism
 - A) Two criteria for a correct mechanism
 - 1) The rate determining step(slow step) must agree with the experimental rate law expression. **CAUTION** Watch out for equilibrium substitution into rate determining step.
 - 2) The elementary steps in the mechanism must add up to be equal to the overall reaction.
 - B) Identifying a catalyst in a reaction mechanism
 - 1) A catalyst will be a reactant in the first step then be consumed in the reaction. A second reaction in the mechanism will produce the catalyst as a product. When adding the elementary steps the catalyst will cancel out and not appear in overall equation even though it was there at the start.
 - C) Identifying an intermediate in a reaction mechanism.
 - 1) An intermediate is not present at the start but is produced during the reaction and then consumed.
5. Collision theory
 - A) Three postulates of the collision. Use them in explaining rates of reactants.
 1. Must have collisions between particles.
 2. Particles must be above the energy of activation in order to react.
 3. Collisions must have the proper orientation(bond angle and molecular shape) and particles above the energy of activation in order to have a successful collision.
6. Factors that affect the rate of a reaction.
 - A) Concentration or partial pressure of the reactants
 1. Rate law expression which can be found by experimentation(comparing rates) or by graphing. In essays explaining what happens to rate, if you have the rate expression use it when you add or remove a substance from the rate law expression.
 - B) Temperature
 1. An increase in temperature produces more particles above the energy of activation and faster rate. You also have more collisions because the particles are traveling faster therefore a greater probability of the proper orientation.
 - C) Catalyst
 1. A catalyst lowers the energy of activation and produces more particles above the energy of activation. You also have a greater probability of the proper orientation.
7. Half-Life
 - A) Only first order half-life is independent of the initial concentration of the substance undergoing decay. From the graph concentration vs time(curve) one can determine if the reaction is first order by looking at the time the substance undergoes two half-life. If the time is the same, then it is first order.

THERMODYNAMICS

1. Heat with temperature change(kinetic energy)
 - A) Specific heat
 1. q = specific heat (mass) (change in temperature)
2. Heat with no temperature change (potential energy)
 - A) Enthalpy (general name is Heat of Reaction)
 1. Heat of formation H°_f elements forming a compound
 2. Heat of Combustion
 3. Heat of Solution (dissolving of a compound in water)
 4. Heat of Neutralization acid-base reaction (titration)
 5. Heat of Fusion solid to liquid
 6. Heat of Sublimation solid to gas
 7. Heat of Vaporization liquid to gas
 8. **5,6 and 7 are phase changes which is an equilibrium system with no change in temperature and the change in free energy is zero.**
 9. The change in enthalpy is positive for an endothermic reaction and negative for an exothermic reaction.
 10. Exothermic reactions are favorable in nature.
 11. Endothermic reactions decrease the temperature of the system because kinetic energy is converted to potential energy therefore $\Delta H > 0$.
 12. Exothermic reactions increase the temperature in the system because potential energy is converted to kinetic energy therefore $\Delta H < 0$.
3. Entropy (S)
 - A) Chaos, randomness
 1. Reactions in nature tend to go to maximum chaos
 2. A positive change in entropy favorable in nature.
 3. $S_{\text{gas}} \gg S_{\text{liquid}} > S_{\text{solid}}$
 4. Predict change in entropy and why? Look at moles of gas on both sides and state order or disorder when answering why.
4. Free Energy (G)
 - A) Energy consumed when a reaction drives to completion.(spontaneous reaction)
5. Spontaneous Reaction
 - A) Criteria
 1. $\Delta G < 0$
 2. Use $\Delta G = \Delta H - T\Delta S$ to explain spontaneity of a chemical reaction and how it is temperature controlled.
 3. If the chemical reaction is a redox reaction, for a spontaneous reaction $\Delta G < 0$, $E > 0$ and equilibrium constant, K , is very large.
 4. If the chemical reaction is not a redox reaction, for a spontaneous reaction $\Delta G < 0$ and equilibrium constant, K , is very large.
6. Thermochemical Equation
 - A) Must have states of matter and change in enthalpy(H) with correct sign.

ACID-BASE

1. Definition
 - A) **Bronsted-Lowry**
 1. An acid is a proton donor(H^+) and a base is a proton acceptor.(central atom of base must have a lone pair of electrons)
 - B) **Lewis**
 1. A Lewis acid is an electron pair acceptor(think how H^+ needs a pair of electrons) and a Lewis base is an electron pair donor.(same definition as Bronsted-Lowry, in order to accept a H^+ must have lone pair on central atom)

2. To be a Lewis acid the central atom has to have an **empty valence orbital**. Group 13 as a central atom will have an empty orbital therefore an acid. (BF₃) The formation of a complex ion is a Lewis acid-base reaction with the transition ion being the acid.

2. Titration

A) **Strong Acid vs Strong Base**

1. At the equivalence point the pH is 7 because the salt of a strong acid and base can't undergo hydrolysis. Neither ion can stress the water equilibrium.
2. No equilibrium involved in this titration. The strong acid and base have dissociated 100%.
3. Up to the equivalence point there is a limiting reactant. It will be the substance coming from the buret. To find pH at any point before the equivalence point, calculate number moles of left over substance (H₃O⁺ or OH⁻) and divide by total volume. (dilution)

B) **Strong Acid (buret) vs Weak Base**

1. At the equivalence point the pH is less than 7 (acidic) and the conjugate acid of the weak base (positive ion) will react with water to form the weak base and hydronium ions. (H₃O⁺)
2. During the titration the strong acid (H₃O⁺) will react with the weak base and produce the conjugate acid of the weak base (positive ion) and water. During this titration you are producing a buffer solution. (weak base and conjugate acid, positive ion)
3. Before the equivalence point must do two calculations to find the pH. The first calculation is a limiting reactant problem with the strong acid being the limiting reactant. This reaction is the formation of the buffer. $\text{H}_3\text{O}^+ + \text{weak base} \rightarrow \text{conjugate acid} + \text{H}_2\text{O}$ (work this problem using moles)
The second reaction is the weak base equilibrium reaction with a common ion (conjugate acid, positive ion)
 $\text{Weak base} + \text{H}_2\text{O} \rightarrow \text{conjugate acid} + \text{OH}^-$ Take the moles from previous problem and convert to molarity using total volume. Both weak base and conjugate acid will have a numerical value at the **start** and OH⁻ will be **zero**.
4. At the equivalence point the strong acid and weak base are limiting reactants and the remaining solution contains a salt of a strong acid and weak base. (positive ion from weak base and negative ion from strong acid) $\text{Positive ion from weak base} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{weak base}$ This is hydrolysis and the concentration of H₃O⁺ and weak base is **zero** at the start. The concentration of the positive ion is moles of weak base at **start** of reaction divided by the total volume (volume of base at start of titration + volume of strong acid needed to reach the equivalence point) The equilibrium constant, $K_a = K_w / K_b$ (for the weak base)
5. Beyond the equivalence point we have excess strong acid. Find the moles of excess H₃O⁺ divide by **total** volume and find pH.

C) **Strong Base (buret) vs Weak Acid**

1. At the equivalence point the pH is greater than 7 (basic) and the conjugate base of the weak acid (negative ion) will react with water to form the weak acid and hydroxide ions.
2. During the titration the strong base (OH⁻) will react with the weak acid and produce the conjugate base of the weak acid (negative ion) and water. During this titration you are producing a buffer solution. (weak acid and conjugate base, negative ion)
3. Before the equivalence point must do two calculations to find the pH. The first calculation is a limiting reactant problem with the strong base being the limiting reactant. This reaction is the formation of a buffer. $\text{OH}^- + \text{weak acid} \rightarrow \text{conjugate base} + \text{H}_2\text{O}$ (work this problem using moles)
The second reaction is the weak acid equilibrium reaction with a common ion (conjugate base, negative ion)
 $\text{Weak acid} + \text{H}_2\text{O} \rightarrow \text{conjugate base} + \text{H}_3\text{O}^+$ Take the moles from previous problem and convert to molarity using total volume. Both weak acid and conjugate base will have a numerical value at the start and H₃O⁺ will be zero. Find pH using H₃O⁺.
4. At the equivalence point the strong base and weak acid are limiting reactants and the remaining solution contains a salt of a weak acid and strong base. (positive ion from strong base and negative ion from weak acid) $\text{Negative ion from weak acid} + \text{H}_2\text{O} \rightarrow \text{OH}^- + \text{weak acid}$ This is hydrolysis and the concentration of OH⁻ and weak acid is **zero** at the start. The concentration of the negative ion is moles of weak acid at **start** of reaction divided by the total

volume(volume of acid at start of titration + volume of strong base needed to reach the equivalence point) The equilibrium constant, $K_b = K_w / K_a$ (for the weak acid)
 5. Beyond the equivalence point we have excess strong base. Find the moles of excess OH^- divide by **total** volume and find pOH then pH.

D) Weak Acid and Weak Base

1. At the equivalence point both ions of the salt will undergo hydrolysis. Can write two hydrolysis reactions and solve for both equilibrium constants. The larger equilibrium constant will determine the pH.
2. The K_a for acetic acid and the K_b for ammonia are the same therefore the pH at the equivalence point would be 7. Very unusual. Most of the time K_a and K_b are not equal to each other.

3. **Titration Errors** Two solutions $M_a V_a = M_b V_b$ One solution and one solid $M_b V_b = \text{mass/molar mass}$

A) All titration errors can be explained by the volume of the buret.

1. Improper washing of the buret. Dilution of solution in buret results in a high buret volume reading.
2. Air bubble in the buret results in a high buret volume reading.
3. Improper washing of the pipet. Dilution of substance being titrated results in low buret volume reading.
4. Impure solid sample will result in a low buret volume reading.
5. Wrong indicator results in a high or low buret volume reading depending on the type of hydrolysis reaction at the equivalence point.

4. Calculations

- A) Dilution ($MV = MV$)
- B) Volume needed to reach equivalence point ($MV = MV$ or $MV = \text{mass/molar mass}$) **Watch out for diprotic acid or base.**
- C) Percent dissociation or ionization
- D) Common ion effect
- E) pH at equivalence point (hydrolysis)
- F) pH at any point up to equivalence point
- G) pH beyond equivalence point (strong acid or base, no equilibrium)
- H) Given % dissociation find equilibrium constant
- I) For a weak acid one-half to equivalence point the pH is equal to the $\text{p}K_a$ of the weak acid.
- J) For a weak base one-half to equivalence point the pOH is equal to the $\text{p}K_b$ of the weak base.

5. Indicator (weak organic acid)

- A) Pick an indicator where the $\text{p}K_a$ of the indicator is equal to the pH at the equivalence point. The pH at the equivalence point is the result of the type of acid and base being titrated (hydrolysis)
- B) The acid part of the indicator has one color and the conjugate base of the acid has a second color. When the concentration of the weak acid is approximately equal to the conjugate base (10:1 ratio maximum), you get the transition color.

6. Diprotic Acid

- A) If $K_{a1} > K_{a2}$ by 10^3 or greater, then the hydronium ion produced in the second dissociation is so small that it is insignificant. (drop) Find pH by using the hydronium ion from the first step.
- B) If $K_{a1} > K_{a2}$ by 10^3 or greater, then the double negative ion of the diprotic acid is equal to K_{a2} .
Example $\text{H}_2\text{C}_2\text{O}_4$ $[\text{C}_2\text{O}_4^{2-}] = K_{a2}$
- C) If the hydronium ion is a common ion for the dissociation of diprotic acid then you can use the summation of the two dissociation reactions. Example: $\text{H}_2\text{C}_2\text{O}_4 + 2 \text{H}_2\text{O} \rightleftharpoons 2 \text{H}_3\text{O}^+ + \text{C}_2\text{O}_4^{2-}$
and $K = K_1 \times K_2$

K_p and K_c

1. The ratio of the rate constant is equal to K_c or K_p
A) $k(\text{forward}) / k(\text{reverse}) = K$
2. Problem Solving
A) Finding K given an equilibrium value of a substance and initial value of a substance.
B) Finding K given the % dissociation.

C) Finding K_p given K_c $K_p = K_c(RT)^{\text{change in gas moles}}$ change in gas moles = moles gas products – moles gas reactants

D) Given K and finding equilibrium values.

3. LeChatelier's Principle

A) Stresses

1. Adding or taking away a substance

2. Temperature change This stress will make the equilibrium constant change. This is the only stress that will do it. Increase temperature the endothermic reaction will be favored. Decrease temperature will favor the exothermic reaction.

3. Change in volume of the container. This stress will only happen if there is at least one gas in equilibrium.

4. Add a solid No change Solids are not in equilibrium expression. Also true for a pure liquid.

5. Add a catalyst No change

6. Add an inert gas No change

7. Add an inert gas under constant pressure means you have made the volume of the container larger therefore use volume of container stress.

K_{sp}

1. Molar Solubility (s)

A) moles per liter Use this unit in the equilibrium expression

B) Watch out for relationships that are not 1:1 Example Ag_3PO_4 $K_{sp} = (s)(2s)^3$

2. Calculations

A) Finding molar solubility given K_{sp}

B) Finding molar solubility with a common ion

C) Finding K_{sp} given solubility. Must use mole/liter finding K_{sp}

D) Will a precipitate form when two solutions are mixed. Find Q and compare to K_{sp} . Remember in calculating the molarities of the ions take into consideration the dilution effect.

E) Placing an ion into two different solutions and which compound precipitates first. Find % of ion that precipitated first when the second ion starts to precipitate. Selected precipitation.

ORGANIC CHEMISTRY

1. Alkane $\text{C}_n\text{H}_{2n+2}$ all sigma bonds (saturated)

A) Methane CH_4

B) Ethane CH_3CH_3

C) Propane $\text{CH}_3\text{CH}_2\text{CH}_3$

D) Butane $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3$

E) Pentane $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$

2. Alkene C_nH_{2n} one pi bond

A) Ethene CH_2CH_2

B) Propene CH_2CHCH_3

C) Butene $\text{CH}_3\text{CH}_2\text{CHCH}_2$

D) Pentene $\text{CH}_3\text{CH}_2\text{CH}_2\text{CHCH}_2$

3. Alkyne $\text{C}_n\text{H}_{2n-2}$ one triple bond (two pi bonds)

A) Ethyne CHCH

B) Propyne CHCCH_3

C) Butyne $\text{CH}_3\text{CH}_2\text{CCH}$

D) Pentyne $\text{CH}_3\text{CH}_2\text{CH}_2\text{CCH}$

4. Isomers

A) Same formula, different structure

5. Functional Groups (where chemical reaction occurs)

A) Double and Triple bonds

B) Organic acid $-\text{COOH}$ CH_3COOH ends in oic

C) Organic base $-\text{NH}_2$ CH_3NH_2 ends in amine

D) Organic alcohol $-\text{OH}$ $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$ ends in ol

E) Aldehyde $\text{R}-\text{CHO}$ CH_3CHO ends in al

F) Ketone $\text{R}-\text{CO}-\text{R}'$ $\text{CH}_3\text{COCH}_2\text{CH}_3$ ends in one